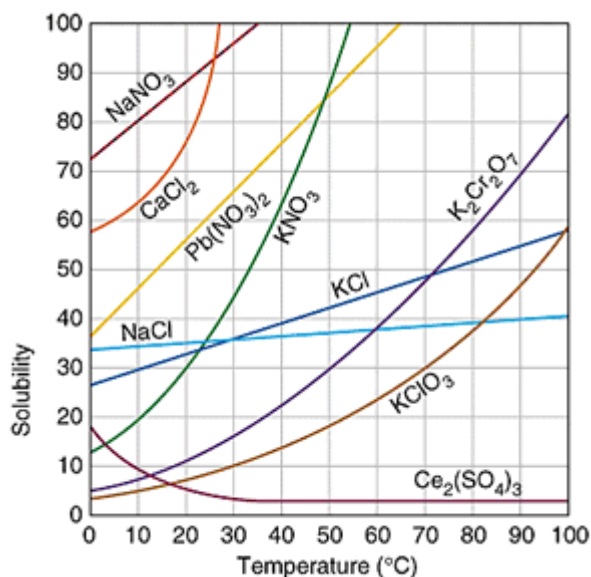


UNIT 8

SOLUBILITY AND PRECIPITATION REACTIONS

Teacher Version



Source: www.dynamicscience.com.au

Contents

- Solubility and Solubility Curves
- Solubility and Precipitation Reactions
- Qualitative Analysis Part 3
- Hard and Soft Water

Key words: aqueous solution, solvent, solute, concentration, molarity, strong electrolyte, weak electrolyte, soluble, solubility, saturated solution, sparingly soluble, insoluble, solubility curve, crystallisation, recrystallisation, precipitate, precipitation, hard water, soft water, limescale, sum, lather, ion exchange

Units which must be completed before this unit can be attempted:

- Unit 1 – Atoms and the Periodic Table
- Unit 2 – Particles, Bonding and Structures
- Unit 3 – Amount of Substance and Measurement
- Unit 4 – Introduction to Physical Chemistry
- Unit 5 – Acids, Bases and Salts
- Unit 6 – Redox Reactions

Estimated Teaching Time: 8 hours

Lesson 1 – What is solubility and what are saturated solutions?

a) Solubility and Solubility Curves

(i) Solubility and Saturated Solutions



Summary Activity 1.1: What is a solution?

- What do the terms “solution”, “solute” and “solvent” mean?
- What is the most common solvent? What is the name given to solutions containing this solvent?
- Can you give an example of a solution?
- What is molarity?
- What is an electrolyte? What is the difference between a strong and a weak electrolyte?

- Solution: mixture of two components, evenly distributed and in the same phase (usually liquid), solute = minor component of a solution; solvent = major component of a solution
- Water; aqueous solutions
- Sea water, brine, limewater
- Moles of solute per cubic decimetre of solution
- A substance which can form free ions in solution; strong electrolytes completely dissociate into ions in solution, weak electrolytes only partially dissociate

- In SS1 (Unit 2 – Particles, Bonding and Structure) you learned that solutions consisted of a solute dissolved in a solvent
- **Solubility** is the ability of a solute to dissolve in a solvent; the solubility of a solute in a solvent is defined as the maximum possible **concentration** of the solute in that solvent; in this course we will consider only **aqueous solutions** (ie solutions in which the solvent is water)
- There are different ways to measure concentration, but the one we use in this course is the moles of solute per cubic decimetre of solution (mol dm^{-3}); this is also known as **molarity**
- The solubility of solutes varies with temperature, so when describing the solubility of a solute it is important to state the temperature; generally, the solubility of solids and liquids increases with increasing temperature but the solubility of gases decreases with increasing temperature
Eg The solubility of sodium chloride (NaCl) in water at 25 °C is 6.2 mol dm^{-3} .
 The solubility of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in water is 5.1 mol dm^{-3} at 25 °C and 14 mol dm^{-3} at 50 °C
 The solubility of ammonia (NH_3) in water is 32 mol dm^{-3} at 20 °C and 17 mol dm^{-3} at 50 °C
- If the maximum possible molarity of a solution is low (ie below 0.2 mol dm^{-3}) the solute is described as slightly soluble; if the solubility is much lower than this the solute might be described as sparingly soluble or insoluble
- A solution which is at its maximum possible concentration at that temperature is said to be **saturated**; it is not possible to dissolve any more solute in a solution which is saturated; the solubility of a substance in water can also therefore be defined as the concentration of a saturated solution; if a solution is not saturated, its concentration is less than the maximum possible and it should be possible to dissolve more of the solute in it

Unit 8 – Solubility and Precipitation Reactions

The solubility of a substance in a solvent is the concentration of a saturated solution of that substance in the solvent

A saturated solution is a solution which is at its maximum possible concentration

- Solubility in water is generally restricted to ionic compounds and molecular substances; metals and giant structures will not dissolve in water unless there is a chemical reaction with the water
 - Ionic compounds, if soluble, are strong electrolytes, which means that they dissociate completely to form ions in solution:
eg $\text{NaCl(s)} \rightarrow \text{Na}^{\text{+}}(\text{aq}) + \text{Cl}^{\text{-}}(\text{aq})$
 - Most molecular substances, if soluble, are usually weak electrolytes or non-electrolytes; weak electrolytes only dissociate slightly to form ions in solution:
eg $\text{NH}_3(\text{g}) \rightarrow \text{NH}_3(\text{g}) \rightleftharpoons \text{NH}_4^{\text{+}}(\text{aq}) + \text{OH}^{\text{-}}(\text{aq})$
 - non-electrolytes do not dissociate at all:
eg $\text{C}_2\text{H}_6\text{O(l)} \rightarrow \text{C}_2\text{H}_6\text{O(aq)}$
 - a very small number of molecular substances are strong electrolytes:
eg $\text{HCl(g)} \rightarrow \text{H}^{\text{+}}(\text{aq}) + \text{Cl}^{\text{-}}(\text{aq})$
- The process of dissolving a solute in a solvent is reversible; it is possible for the solute to separate itself from the solution and form the pure substance (usually as a solid, in which case it settles at the bottom of the container, or as a gas, in which case it escapes); a saturated solution is in a state of **dynamic equilibrium** (the solute is dissolving in the solvent at the same rate as it is separating it out from the solvent)



Test your knowledge 1.2: Solubility and Saturated Solutions

At 50 °C, it is possible to dissolve up to 244 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in 100 cm^3 of water.

- What is the solubility of glucose in water (in mol dm^{-3}) at 50 °C?
- What is the maximum mass of glucose which can dissolve in 25 cm^3 of water at 50 °C?
- A student attempts to dissolve 135 g of glucose in 60 cm^3 of water at 50 °C. Will she succeed? Explain your answer.
- Suggest two ways in which glucose could be crystallised from a solution.
A saturated solution of calcium hydroxide (Ca(OH)_2) is known as limewater. At 25 °C, limewater has a molarity of 0.020 mol dm^{-3} .
- Write an equation, using state symbols for the equilibrium which exists in limewater
- Calculate the minimum mass of calcium hydroxide required to prepare 250 cm^3 of limewater
- A student attempts to dissolve 2.0 g of calcium hydroxide in 100 cm^3 of water at 25 °C. Will he succeed? Explain your answer
- Which is more soluble in water? Glucose or calcium hydroxide? Explain.
- Are glucose and calcium hydroxide electrolytes? Explain your answer.

Unit 8 – Solubility and Precipitation Reactions

- (a) 13.6 mol dm^{-3}
- (b) 61.2 g
- (c) Yes, because the molarity would be 12.5 mol dm^{-3} which is less than a saturated solution
- (d) By heating the solution until some of the water evaporates, or by cooling the solution
- (e) $\text{Ca(OH)}_2(\text{s}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq})$
- (f) 0.37 g
- (g) No, because the molarity would be 0.27 mol dm^{-3} which is greater than the solubility of Ca(OH)_2
- (h) Glucose is more soluble because its saturated solution has a higher molarity than a saturated solution of calcium hydroxide
- (i) Calcium hydroxide is an electrolyte because when it dissolves it dissociates into its ions; glucose it not an electrolyte because it remains as molecules when it dissolves

Lesson 2 – What is crystallisation and what are solubility curves?

(ii) Crystallisation



Summary Activity 2.1: Preparing salts

- When you prepared salts in SS2 (Unit 5 – Acids, Bases and Salts and Unit 6 – Redox Reactions), how did you separate the pure salt from the solution? How does this technique work?
-

- We heated the salt solution gently until most of the water had evaporated off
- Some of the salt crystallises out during heating because the concentration of the solution increases as the water is removed
- Most of the salt crystallises out during cooling; the water continues to evaporate so the concentration increases, and the solubility decreases as the solution cools down

- Salts and other solid solutes can usually be separated from their solutions by **crystallisation**:
 - first heat the solution to evaporate off some of the water; as the water is removed the concentration of the solution increases; when the concentration becomes equal to the solubility of the solute (ie the solution becomes saturated) the solute will start to crystallise
 - then allow the solution to cool; the water will continue to evaporate, causing more solute to crystallise; also as the solution cools down the solubility decreases (most solids become less soluble at lower temperatures) so more solute crystallises
 - don't allow all the solvent to evaporate during heating; this will cause the temperature to rise rapidly and this may cause the solute to burn or decompose
- The technique of crystallisation (or recrystallisation) can also be used as a **purification** technique:
 - the pure solid should dissolve in hot water but crystallise in cold water
 - some impurities will not dissolve in water and so they can be filtered out after the solid dissolves in hot water
 - other impurities dissolve even in cold water and remain in solution after the pure solid crystallises

Unit 8 – Solubility and Precipitation Reactions



Practical 2.2: Purify a sample of copper sulphate by recrystallisation

- 1) Pour 30 cm³ of distilled water into a 250 cm³ beaker
- 2) Add impure copper sulphate to the beaker using a spatula; stir the mixture gently so that the solid dissolves; keep adding the impure copper sulphate until it stops dissolving (you might need 10 g of copper sulphate or even slightly more)
- 3) Add 3 cm³ of 1 mol dm⁻³ H₂SO₄ and stir the mixture gently
- 4) Place some filter paper in a funnel; place the funnel over an evaporating dish and pour the copper sulphate solution into the funnel
- 5) Now place the evaporating dish containing the filtrate onto a gauze on a tripod above a Bunsen burner (or onto a sand bath); heat the solution until just over half of the liquid has evaporated
- 6) Allow the mixture to cool (it's best to leave it for a few days)
- 7) Carefully decant the excess solution and use a clean spatula to scrape any crystals onto a piece of filter paper
- 8) Transfer the crystals onto another piece of filter paper; fold the filter paper over the crystals and squeeze them gently

At what point during this procedure are the insoluble impurities removed?

At what point during this procedure are the soluble impurities removed?

Equipment needed per group: 15 g of hydrated copper sulphate, spatula, 50 cm³ measuring cylinder; stirring rod, 250 cm³ beaker, 3 pieces of filter paper, funnel, evaporating dish, tripod, gauze and Bunsen burner OR sand bath, access to distilled water

- The insoluble impurities are removed when the solid is filtered

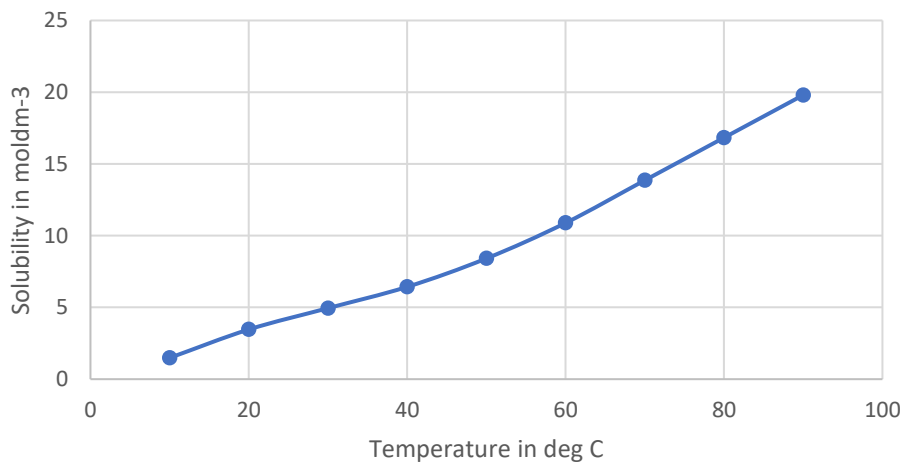
- The soluble impurities are removed when the solid is decanted

Can't do this experiment? Watch this video instead: <https://www.youtube.com/watch?v=SAU-gptAFe0>

(iii) Solubility Curves

- Most solids are more soluble in water at high temperatures than at low temperatures; it is useful to know the solubility of different solids at different temperatures and **solubility curves** are used for this purpose; the solubility curve for KNO₃ in water is shown below:

Solubility curve for KNO₃ in water



Unit 8 – Solubility and Precipitation Reactions



Test your knowledge 2.3: Using solubility curves

- What is the solubility of KNO_3 at 55°C ?
- What is the molarity of a saturated solution of KNO_3 at 15°C ?
- What mass of KNO_3 can be dissolved in 50 cm^3 of water at 25°C ?
- What volume of water is required to dissolve 20 g of KNO_3 at 70°C ?
- 10 g of KNO_3 is dissolved in 10 cm^3 of boiling water at 100°C and the solution is allowed to cool. At what temperature will crystals start to form?
- What is the minimum temperature required to dissolve 15 g of KNO_3 in 20 cm^3 of water?
- Is it possible to dissolve completely 30 g of KNO_3 in 25 cm^3 of water at 60°C ?



Extension 2.4: Using solubility curves

Completed Exercise 2.3? Use the solubility curves shown on the first page of this topic to set questions for each other to solve!

- Approx 9 mol dm^{-3}
- Approx 2.5 mol dm^{-3}
- Solubility = approx. 4 mol dm^{-3} so $n = 4 \times 0.05 = 0.2$ so $m = 20\text{ g}$
- Solubility = approx. 14 mol dm^{-3} , $n = 20/101 = 0.2$ so $V = n/C = 0.014\text{ dm}^3$ or 14 cm^3
- $n = 10/101 = 0.1$ so $C = 0.1/0.01 = 10\text{ mol dm}^{-3}$ so $T = 55 - 57^\circ\text{C}$
- $n = 15/101 = 0.15$ so $C = 0.15/0.02 = 7.5\text{ mol dm}^{-3}$ so $T = 46 - 48^\circ\text{C}$
- $n = 30/101 = 0.3$ so $C = 0.3/0.025 = 12\text{ mol dm}^{-3}$ but solubility = 11 mol dm^{-3} so not all will dissolve

Lesson 3 – What is precipitation and what is a precipitation reaction?

b) Solubility and Precipitation Reactions

- In SS1 (Unit 2 – Particles, Bonding and Structure) you learned about why most, but not all, ionic compounds were soluble in water; in SS2 (Topic 5 – Acids, Bases and Salts) you learned about the solubility of a few common bases and salts



Summary Activity 3.1: Solubility of Ionic Compounds

- What causes ionic compounds to dissolve in water?
 - Why are some ionic compounds not soluble in water?
 - Can you name any ionic compounds which are soluble in water?
 - Can you name any ionic compounds which are not soluble in water?
- Ionic compounds dissolve in water because the positive ions are attracted to the electronegative O atom in water and the negative ions are attracted to the electropositive H atom in water
 - The attraction between the ions and water has to be stronger than the attraction of the ions to each other; in some cases the ions are attracted to each other more strongly than they are attracted to water
 - sodium chloride, ammonium sulphate, copper sulphate
 - silver chloride, calcium carbonate etc

Unit 8 – Solubility and Precipitation Reactions

(i) Predicting the solubility of ionic compounds

- Most common ionic compounds are soluble in water (ie have solubilities in water of 0.5 mol dm^{-3} or greater at 25°C); however there are a number of exceptions; by following a few general rules, it is possible to predict the solubility of most ionic compounds in water:
 - Rule 1: All compounds containing the nitrate ion (NO_3^-) are soluble
 - Rule 2: All compounds containing Group 1 ions (eg H^+ , Li^+ , Na^+ , K^+ etc) and ammonium ions (NH_4^+) are soluble
 - Rule 3: Most compounds containing halide ions (Cl^- , Br^- , I^-) are soluble except the halides of lead (eg PbCl_2) and silver (eg AgBr)
 - Rule 4: Most compounds containing sulphate ions (SO_4^{2-}) are soluble except the sulphates of lead, calcium, strontium and barium (PbSO_4 , CaSO_4 , SrSO_4 , BaSO_4)
 - Rule 5: Most compounds containing hydroxide ions (OH^-) and carbonate ions (CO_3^{2-}) are insoluble except those mentioned in Rule 2 (Rule 5); $\text{Ba}(\text{OH})_2$ and $\text{Sr}(\text{OH})_2$ are also soluble



Test your knowledge 3.2: Predicting the Solubility of Ionic Compounds

Give the formula of each of the following ionic compounds, and predict which will be soluble in water:

- | | |
|--------------------------|---------------------------|
| (a) magnesium nitrate | (i) barium carbonate |
| (b) sodium sulphate | (j) potassium carbonate |
| (c) copper (II) chloride | (k) calcium carbonate |
| (d) silver chloride | (l) copper (II) hydroxide |
| (e) lead (II) bromide | (m) lithium hydroxide |
| (f) copper (II) sulphate | (n) barium hydroxide |
| (g) barium sulphate | (o) magnesium hydroxide |
| (h) magnesium sulphate | |



Extension 3.3: Predicting the Solubility of Ionic Compounds

Completed Exercise 3.2? Use the rules to predict the solubility of some more ionic compounds and test each other!

- | | |
|---|---|
| (a) $\text{Mg}(\text{NO}_3)_2$; soluble (Rule 1) | (i) BaCO_3 ; insoluble (Rule 5) |
| (b) Na_2SO_4 ; soluble (Rule 2) | (j) K_2CO_3 ; soluble (Rule 2) |
| (c) CuCl_2 ; soluble (Rule 3) | (k) CaCO_3 ; insoluble (Rule 5) |
| (d) AgCl ; insoluble (Rule 3) | (l) $\text{Cu}(\text{OH})_2$; insoluble (Rule 5) |
| (e) PbBr_2 ; insoluble (Rule 3) | (m) LiOH ; soluble (Rule 2) |
| (f) CuSO_4 ; soluble (Rule 4) | (n) $\text{Ba}(\text{OH})_2$; soluble (Rule 5) |
| (g) BaSO_4 ; insoluble (Rule 4) | (o) $\text{Mg}(\text{OH})_2$; insoluble (Rule 5) |
| (h) MgSO_4 ; soluble (Rule 4) | |

(ii) Precipitation

- If one aqueous solution of an ionic compound is added to another, the cations in the first solution will come in to contact with the anions in the second solution, and the anions in the first solution will come into contact the cations in the second solution; in other words, two new ionic compounds will be created
Eg if $\text{NaCl}(\text{aq})$ and $\text{AgNO}_3(\text{aq})$ are mixed together, the Na^+ will come into contact with the NO_3^- (NaNO_3) and the Ag^+ will come into contact with the Cl^- (AgCl)

Unit 8 – Solubility and Precipitation Reactions

- If both of these new compounds are soluble, the ions all mix together freely and no chemical reaction takes place; if either of the new compounds is insoluble, however, the two ions will join together to form a solid; this solid is called a **precipitate**; the formation of an insoluble ionic compound by mixing two different solutions together is called **precipitation**

Eg if NaCl(aq) and AgNO₃(aq) are mixed together:

- NaNO₃ and AgCl can be formed
- NaNO₃ is soluble so the Na⁺ and NO₃⁻ ions remain in solution
- AgCl, is insoluble so the ions join together and a precipitate is observed
- the ionic equation for the reaction is **Ag⁺(aq) + Cl⁻(aq) → AgCl(s)**



Test your knowledge 3.4: Predicting Precipitation

Predict whether or not a precipitate will form when the following solutions are mixed. If precipitation occurs, write the ionic equation for the reaction taking place:

- barium chloride solution and sodium sulphate solution
- lead nitrate solution and sodium chloride solution
- copper sulphate solution and sodium hydroxide solution
- sodium carbonate solution and ammonium nitrate solution
- potassium carbonate solution and calcium chloride solution
- magnesium chloride solution and sulphuric acid
- silver nitrate solution and hydrochloric acid
- sodium chloride solution and nitric acid



Extension 3.5: Predicting Precipitation

Completed Exercise 3.4? Use the rules to predict some other pairs of solutions which will form a precipitate when mixed together, and write ionic equations for the reactions taking place!

- precipitate: $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$
- precipitate: $\text{Pb}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s})$
- precipitate: $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^{-}(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$
- no precipitate
- precipitate: $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$
- no precipitate
- precipitate: $\text{Ag}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq}) \rightarrow \text{AgCl}(\text{s})$
- no precipitate

Lesson 4 – How can we prepare insoluble salts?



Practical 4.1: Observe precipitation reactions

You are provided with six bottles containing the following solutions:

- A is silver nitrate (AgNO₃)
- B is hydrochloric acid (HCl)
- C is sulphuric acid (H₂SO₄)
- D is barium chloride (BaCl₂)
- E is copper sulphate (CuSO₄)

Unit 8 – Solubility and Precipitation Reactions

F is sodium hydroxide (NaOH)

Each bottle should come with its own 10 cm³ measuring cylinder

- 1) Pour 5 cm³ of solution A into a test tube
- 2) Add 5 cm³ of solution B to the same test tube; note whether you see a precipitate or not
- 3) Write “precipitate” or “no precipitate” in the correct box of the table below
- 4) Repeat steps 2 – 3 using solutions C, D, E and F instead of B
- 5) Continue mixing the solutions until each combination has been mixed together once

Write ionic equations for any reactions taking place

	A	B	C	D	E	F
A						
B						
C						
D						
E						

Equipment needed per group: access to labelled bottles containing 0.1 mol dm⁻³ AgNO₃, HCl, H₂SO₄, BaCl₂, CuSO₄ and NaOH and one 10 cm³ measuring cylinder for each bottle, 15 x test tubes

Expected observations:

	A	B	C	D	E	F
A		precipitate	precipitate	precipitate	precipitate	precipitate
B			no precipitate	no precipitate	no precipitate	no precipitate
C				precipitate	no precipitate	no precipitate
D					precipitate	no precipitate
E						precipitate

Equations:

AB: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$; AC: $2\text{Ag}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{Ag}_2\text{SO}_4(\text{s})$; AD: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$;
 AE: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$; AF: $\text{Ag}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{AgOH}(\text{s})$; CD: $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$;
 DE: $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$; EF: $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{Cu}(\text{OH})_2(\text{s})$

(iii) Preparing Insoluble Salts



Summary Activity 4.2: Preparation of soluble salts

In SS2 (Topic 5 – Acids, Bases and Salts and Topic 6 – Redox Reactions) you prepared some soluble salts.

- Which salts did you prepare and how did you prepare them?
- What types of reaction were used to prepare each salt?
- How was the salt separated from the rest of the reaction mixture?
- Which of the reactants should be in excess?

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- Copper sulphate was prepared by reacting excess copper oxide with sulphuric acid; this is an acid-base (or neutralisation) reaction; the excess copper oxide (which is insoluble) was removed by filtration and the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool)
$$\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$$
- Ammonium sulphate was prepared by reacting ammonia with sulphuric acid in a 2:1 ratio; this is an acid-base (or neutralisation) reaction; both reactants are also soluble so the exact quantities were needed; the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool)
$$2\text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4$$
- Zinc sulphate was prepared by reacting excess zinc with dilute sulphuric acid; this is a redox reaction; the excess zinc was removed by filtration and the soluble salt was extracted by crystallisation (the salt solution was heated and then allowed to cool)
$$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$$

- Soluble salts can be prepared by neutralisation reactions or by redox reactions; you learned about the preparation of soluble salts in SS2 (Topic 5 - Acids, Bases and Salts and Topic 6 – Redox Reactions); now you will learn about the preparation of **insoluble** salts
- Insoluble salts can be prepared very easily by precipitation reactions; if a solution containing the anion is added to a solution containing the cation, a precipitate of the insoluble salt is formed; a pure sample of the insoluble salt can be separated from the rest of the mixture by filtration
Eg the addition of lead nitrate solution to sodium chloride solution will result in the precipitation of the insoluble salt lead chloride



Practical 4.3: Prepare a sample of the insoluble salt lead chloride

(caution – lead compounds are toxic; wash your hands thoroughly after carrying out this experiment and do not pour the lead (II) nitrate solution down the sink)

- 6) Pour 5 cm³ of 1 mol dm⁻³ lead (II) nitrate solution into a test tube
- 7) Pour 5 cm³ of 2 mol dm⁻³ sodium chloride solution into the same test tube
- 8) Place a stopper on the test tube and shake the contents thoroughly
- 9) Place a piece of filter paper into a funnel and pour the contents of the test tube into the funnel; allow the filtrate to run into a beaker
- 10) When the filtration is complete, pour a small quantity of cold distilled water into the funnel
- 11) Use a spatula to scrape the residue off the filter paper and onto another dry piece of filter paper
- 12) Fold the filter paper of the solid sample and squeeze gently

Write an ionic equation for the reaction taking place during this preparation.

What type of reaction is taking place?

Does it matter which reactant is in excess? Explain your answer.

Equipment needed per group: 5 cm³ of 1 mol dm⁻³ lead (II) nitrate solution, 5 cm³ of 2 mol dm⁻³ sodium chloride solution; 2 x 10 cm³ measuring cylinders test tube with stopper; funnel, 2 x filter paper, small beaker (50 cm³), spatula



- **Precipitation**

- **It doesn't matter; whichever reactant is in excess will be removed during the filtration process**

Can't do this experiment? Watch this video instead: <https://www.youtube.com/watch?v=E1ODnGe9LnM>

Lesson 5 – How can we use precipitation reactions to identify cations in solution?

c) Qualitative Analysis Part 3 – Precipitation Reactions

- Qualitative analysis is the experimental identification of a substance or of a particular species present in a substance; you learned about some simple qualitative tests for cations, anions and gases in Unit 5 – Acids, Bases and Salts and Unit 6 – Redox Reactions



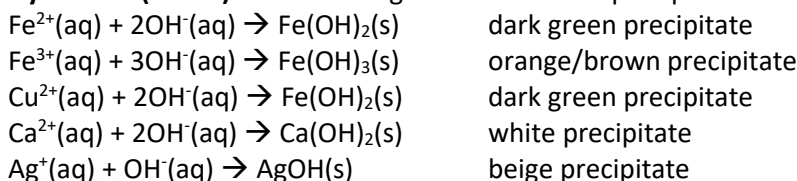
Summary Activity 5.1: Qualitative Analysis of Cations

- Describe two different tests to confirm the presence of H^+ ions in solution
- Describe a test to confirm the presence of NH_4^+ ions in solution

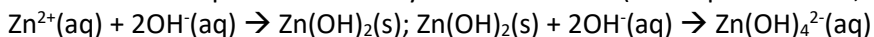
- Add blue litmus paper; it will turn red; add a sample of calcium carbonate; a gas will be given off which turns limewater milky
- Add sodium hydroxide solution and warm; a pungent gas should be given off which turns red litmus paper blue

- Precipitation reactions are very useful in qualitative analysis – they are commonly used to identify particular cations and anions in solution:

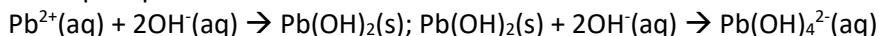
- Sodium hydroxide (NaOH)** solution can give characteristic precipitates with many common metal ions:



In some cases, the hydroxide precipitates formed are amphoteric – this means that they react with alkali as well as acid, and dissolve in aqueous sodium hydroxide solution (see Topic 5 – Acids, Bases and Salts)



white precipitate which dissolves in excess NaOH



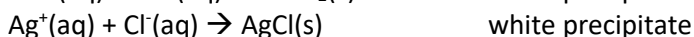
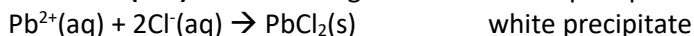
white precipitate which dissolves in excess NaOH



white precipitate which dissolves in excess NaOH

A small number of cations (NH_4^+ , Na^+ , K^+ , Ba^{2+}) do not give a precipitate with hydroxide ions

- Hydrochloric acid (HCl)** solution can give characteristic precipitates with a small number of metal ions:



Most other cations do not give a precipitate with chloride ions; this reaction allows Pb^{2+} to be distinguished from Zn^{2+} and Al^{3+}

Unit 8 – Solubility and Precipitation Reactions

- **Summary: precipitation tests for cations**

Cation	Test	Observation
Fe ²⁺	Add NaOH	Dark green precipitate
Fe ³⁺	Add NaOH	Orange brown precipitate
Cu ²⁺	Add NaOH	Pale blue precipitate
Ca ²⁺	Add NaOH	White precipitate insoluble in excess NaOH
Zn ²⁺ or Al ³⁺	Add NaOH Add HCl	White precipitate soluble in excess NaOH No precipitate
Pb ²⁺	Add NaOH Add HCl	White precipitate soluble in excess NaOH White precipitate



Practical 5.2: Qualitative Analysis Part 3a - use precipitation reactions to identify cations in solution

You are given eight solutions, labelled A, B, C, D, E, F, G and H; each solution contains one cation from the following: Pb²⁺, Ca²⁺, Fe²⁺, Fe³⁺, Zn²⁺, Cu²⁺, Al³⁺ and Na⁺

- Pour 2 cm³ of solution A into a test tube
- Add a few drops of dilute NaOH, then add dilute NaOH until the test tube is almost full; note your observations
- Repeat steps (a) to (b) using solutions B – H
- Pour 2 cm³ of solution A into a test tube
- Add a few drops of dilute HCl; record your observations
- Repeat steps (d) to (e) using solutions B – H

Use your observations to identify which cation is in which solution; present your answers in a table similar to the one below:

Solution	Observations			cation present
	few drops NaOH	excess NaOH	few drops HCl	
A				
B				
C				
D				
E				
F				
G				
H				

Which two cations cannot be distinguished by this combination of tests?

Unit 8 – Solubility and Precipitation Reactions

Chemicals needed: minimum 0.1 mol dm⁻³ solutions of: FeSO₄ (labelled A), Na₂CO₃ (labelled B), ZnSO₄ (labelled C), Ca(NO₃)₂ (labelled D), CuSO₄ (labelled E), Pb(NO₃)₂ (labelled F), Al₂(SO₄)₃ (labelled G) and FeCl₃ (labelled H) - around 10 cm³ per group prepared in a single bottle, each with its own dropping pipette; also 0.5 - 1 mol dm⁻³ of the following solutions: NaOH, HCl - up to 20 cm³ per group; each group needs its own bottle with its own dropping pipette

Apparatus needed per group: 16 test tubes, 1 test tube rack

Expected observations:

Solution	Observations			cation present
	few drops NaOH	excess NaOH	few drops HCl	
A	dark green precipitate	no change	no change	Fe ²⁺
B	no change	no change	no change	Na ⁺
C	white precipitate	dissolves – colourless solution	no change	Zn ²⁺ or Al ³⁺
D	white precipitate	no change	no change	Ca ²⁺
E	pale blue precipitate	no change	no change	Cu ²⁺
F	white precipitate	dissolves – colourless solution	white precipitate	Pb ²⁺
G	white precipitate	dissolves – colourless solution	no change	Zn ²⁺ or Al ³⁺
H	orange-brown precipitate	no change	no change	Fe ³⁺

Zn²⁺ and Al³⁺ cannot be distinguished by this combination of tests

Lesson 6 – How can we use precipitation reactions to identify anions in solution?



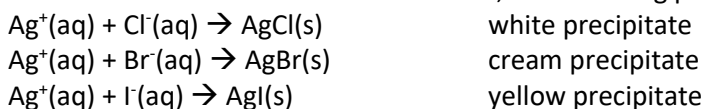
Summary Activity 6.1: Qualitative Analysis of Anions

- Describe two different tests to confirm the presence of OH⁻ ions in solution
- Describe a test to confirm the presence of NO₃⁻ ions in solution
- Describe a test to confirm the presence of CO₃²⁻ ions in solution
- Describe a test to confirm the presence of SO₃²⁻ ions in solution

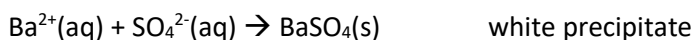
- Add red litmus paper; it will turn blue; add some ammonium chloride and warm; a pungent gas should be given off which turns red litmus paper blue
- Add sodium hydroxide solution and aluminium powder and heat; a pungent gas should be given off which turns red litmus paper blue
- Add HCl(aq); a gas will be given off which turns limewater milky
- Add HCl(aq); a gas will be given off which turns blue litmus paper red and turns dichromate paper green

Unit 8 – Solubility and Precipitation Reactions

- Silver nitrate (AgNO₃) solution** can give characteristic precipitates with most anions; however if these precipitates contain basic ions such as SO₃²⁻, CO₃²⁻ or OH⁻, the precipitates will react with acids; Ag₂SO₃, AgOH and Ag₂CO₃ are all insoluble but react with acids; these precipitates therefore do not form if **dilute nitric acid (HNO₃)** is added before the silver nitrate solution; the following precipitates form even in acidic conditions:



- Barium chloride (BaCl₂) solution** can give characteristic precipitates with SO₃²⁻, SO₄²⁻ and CO₃²⁻, some anions. BaSO₃ and BaCO₃ contain basic ions so react with acids. These precipitates therefore do not form if **dilute hydrochloric acid (HCl)** is added before the barium chloride solution. Under those conditions, only BaSO₄ will form:



- Summary: precipitation tests for anions**

Cation	Test	Observation
Cl ⁻	Add HNO ₃ and then AgNO ₃	white precipitate
Br ⁻	Add HNO ₃ and then AgNO ₃	cream precipitate
I ⁻	Add HNO ₃ and then AgNO ₃	yellow precipitate
SO ₄ ²⁻	Add HCl and then BaCl ₂	white precipitate
SO ₃ ²⁻	Add BaCl ₂ Then add HCl	White precipitate Dissolves in acid; gas evolved which turns blue litmus red and acidified dichromate paper green
CO ₃ ²⁻	Add BaCl ₂ Then add HCl	White precipitate soluble in excess NaOH Dissolves in acid; gas evolved which turns limewater milky

Unit 8 – Solubility and Precipitation Reactions



Practical 6.2: Qualitative Analysis Part 3b - use precipitation reactions to identify anions in solution

You are given six solutions, labelled A, B, D, E, H and I; each solution contains one anion from the following: NO_3^- , I^- , SO_4^{2-} , Cl^- , SO_3^{2-} , CO_3^{2-}

- (a) Pour 2 cm³ of solution A into a test tube
- (b) Add 2 cm³ of HNO_3 , and then a few drops of AgNO_3 ; note your observations
- (c) Repeat steps (e) – (f) using solutions B – H
- (d) Pour 2 cm³ of solution A into a test tube
- (e) Add a few drops of BaCl_2 ; note your observations
- (f) Repeat steps (e) – (f) using solutions B – H
- (g) Pour 2 cm³ of each solution which gave a precipitate with BaCl_2 into separate test tubes
- (h) To each test tube, add 2 cm³ of HCl , and then a few drops of BaCl_2 ; note your observations

Use your observations to identify which anion is in which solution; present your answers in a table similar to the one below:

Solution	Observations			anion present
	HNO_3 and AgNO_3	BaCl_2	HCl and BaCl_2	
A				
B				
D				
E				
H				
I				

Which two anions cannot be distinguished by this combination of tests? What test could be used to distinguish between these two anions?

Chemicals needed: minimum 0.1 mol dm⁻³ solutions of: FeSO_4 (labelled A), Na_2CO_3 (labelled B), KI (labelled D), KNO_3 (labelled E), FeCl_3 (labelled H), Na_2SO_3 (labelled I) - around 10 cm³ per group prepared in a single bottle, each with its own dropping pipette; also 0.05 mol dm⁻³ AgNO_3 , 0.1 mol dm⁻³ BaCl_2 , 1 mol dm⁻³ HCl , 1 mol dm⁻³ HNO_3 - up to 50 cm³ per group; each group needs its own bottle with its own dropping pipette

Apparatus needed per group: 17 test tubes, 1 test tube rack

Expected observations:

Solution	Observations			anion present
	HNO_3 and AgNO_3	BaCl_2	HCl and BaCl_2	
A	white precipitate	white precipitate	white precipitate	SO_4^{2-}
B	no change	white precipitate	no change	CO_3^{2-} or SO_3^{2-}
D	yellow precipitate	no change	-	I^-
E	no change	no change	-	NO_3^-
H	white precipitate	no change	-	Cl^-
I	no change	white precipitate	no change	CO_3^{2-} or SO_3^{2-}

CO_3^{2-} and SO_3^{2-} cannot be distinguished by this combination of tests; they could be distinguished by adding $\text{CaCl}_2(\text{aq})$ and then adding $\text{HCl}(\text{aq})$ to the resulting precipitate; the gas evolved from CO_3^{2-} will turn limewater milky; the gas evolved from SO_3^{2-} will turn blue litmus red and turn dichromate paper green

Unit 8 – Solubility and Precipitation Reactions



Test your knowledge 6.3: Using precipitation to distinguish between different solutions

Describe a simple test to show how you would distinguish between the following solutions:

- (a) sodium chloride and calcium chloride
- (b) zinc nitrate and calcium nitrate
- (c) iron (II) sulphate, copper sulphate and iron (III) sulphate
- (d) zinc nitrate and lead nitrate
- (e) sodium chloride and sodium nitrate
- (f) sodium chloride and sodium sulphate
- (g) sodium sulphate and sodium carbonate

- (a) add NaOH (aq); no reaction with NaCl; white precipitate with $\text{Ca}(\text{OH})_2$
- (b) add NaOH (aq) dropwise and then in excess; white precipitate with $\text{Ca}(\text{NO}_3)_2$ is insoluble in excess NaOH; white precipitate with $\text{Pb}(\text{NO}_3)_2$ dissolves in excess NaOH
- (c) add NaOH (aq); dark green precipitate with FeSO_4 ; pale blue precipitate with CuSO_4 ; orange/brown precipitate with $\text{Fe}_2(\text{SO}_4)_3$
- (d) add HCl (aq); no reaction with $\text{Ca}(\text{NO}_3)_2$; white precipitate with $\text{Pb}(\text{NO}_3)_2$
- (e) Add AgNO_3 (aq); no reaction with NaNO_3 ; white precipitate with NaCl
- (f) Add BaCl_2 (aq); no reaction with NaCl; white precipitate with Na_2SO_4
- (g) Add BaCl_2 (aq) then add HCl(aq); white precipitate with Na_2SO_4 is insoluble in HCl; white precipitate with Na_2CO_3 dissolves in HCl



Extension 6.4: Further qualitative analysis

Completed Exercise 6.3? Describe simple tests to distinguish between the following solutions: (these questions require you to use qualitative analysis from previous units)

- (a) sodium nitrate and nitric acid
 - (b) sodium hydroxide and sodium carbonate
 - (c) sodium nitrate and ammonium nitrate
 - (d) sodium nitrate and water
- (a) add blue litmus paper; it turns red in nitric acid but not in sodium nitrate
 - (b) add HCl; observe bubbles with sodium carbonate but not with sodium hydroxide, or add magnesium chloride solution (or any solution containing a +2 ion); a precipitate forms in both cases; with the carbonate, the precipitate will give off bubbles when it dissolves, but the hydroxide will dissolve without giving off bubbles
 - (c) Add NaOH and heat; with ammonium nitrate, a pungent gas will be given off which turns red litmus blue; with sodium nitrate there will be no reaction
 - (d) Add aluminium powder and sodium hydroxide and heat; with sodium nitrate, a pungent gas will be given off which turns red litmus blue; with water there will be no reaction

Lesson 7 – What is hard water?

d) Hard and Soft Water

(i) definition of hard water

- Hard water is water which contains a high concentration of dissolved Ca^{2+} , Mg^{2+} and Fe^{2+} ions; these ions are commonly found in rocks and can get into the water supply when water comes into contact with these rocks; the presence of these ions in water causes a number of precipitation reactions to take place which can cause problems; soft water is water which contains only a very low concentration of these ions

(ii) disadvantages of hard water

- The first problem caused by the presence of these ions is that when heated, these ions react with dissolved hydrogencarbonate ions (HCO_3^-) in water to form calcium carbonate, which is insoluble:

$$\text{Ca}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$

$$\text{Mg}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow \text{MgCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
 The compound $\text{CaCO}_3(\text{s})$ is known as limescale and can form on the surface of heating elements, especially in kettles, washing machines, dishwashers and industrial water heaters; limescale reduces the efficiency heat transfer and it can also restrict the flow of water if it builds up in pipes
- The second problem caused by the presence of these ions is they react with the anions present in soap, such as $\text{C}_{17}\text{H}_{35}\text{COO}^-$ (stearate ions) to form a precipitate:

$$\text{Ca}^{2+}(\text{aq}) + \text{C}_{17}\text{H}_{35}\text{COO}^-(\text{aq}) \rightarrow \text{Ca}(\text{C}_{17}\text{H}_{35}\text{COO})_2(\text{s})$$
 This precipitate is known as scum and it removes the stearate ions from the water; as a result the lather which should be formed in water by the stearate ions does not form; so soap is less effective in hard water

(iii) advantages of hard water

- Ca^{2+} , Mg^{2+} and Fe^{2+} are important minerals for the human body, and hard water can be a useful source of these minerals; if the concentration of these ions are very high, however, the disadvantages of hard water are judged to outweigh the benefits

(iv) temporary and permanent hardness

- The hardness of water is described as temporary if the main anion present in the water is HCO_3^- ; when heated, HCO_3^- decomposes into CO_3^{2-} which forms a precipitate with Ca^{2+} , Mg^{2+} and Fe^{2+} ions;

$$\text{Ca}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow \text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$

$$\text{Mg}^{2+}(\text{aq}) + 2\text{HCO}_3^-(\text{aq}) \rightarrow \text{MgCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
 This precipitate can be removed by filtration and the resulting water is soft; temporary hardness can therefore be removed simply by heating the water
- The hardness of water is described as permanent if the main anion present in the water is not HCO_3^- ; in such cases it is usually SO_4^{2-} or Cl^- ; these ions do not form precipitation reactions on heating; permanent hardness therefore cannot be removed just by heating the water

Unit 8 – Solubility and Precipitation Reactions

(v) softening water

- Other than simply boiling (which only removes temporary hardness), there are three main techniques available for softening water (ie removing Ca^{2+} , Mg^{2+} and Fe^{2+} ions):
 - Precipitation:** all of the ions responsible for water hardness form insoluble carbonates; the addition of a solution of sodium carbonate (Na_2CO_3) to the water will therefore remove all three of these ions as precipitates:
$$\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$$
$$\text{Mg}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{MgCO}_3(\text{s})$$
$$\text{Fe}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{FeCO}_3(\text{s})$$
The resulting water, when filtered, is soft
 - Distillation:** when water is boiled, any dissolved ions do not boil but remain in the residue (as ionic compounds have a high boiling point); when condensed again, the boiled water has lost its ions and is therefore soft
 - Ion Exchange:** ion exchange resins contain weakly attached sodium ions (Na^+); when hard water is passed through these resins, the Ca^{2+} and Mg^{2+} ions attach more strongly to the resin and replace the Na^+ ions; the Ca^{2+} and Mg^{2+} ions in the water are therefore replaced with Na^+ ions and the water is now soft

(vi) testing the hardness of water

- The degree of hardness in different water samples can be compared simply by adding soap solution gradually until a lather is formed; the greater the quantity of soap solution required, the greater the levels of hardness in the water



Practical 7.1: Test the Hardness of Water

Collect four different samples of water from four different sources (such as tap water, rainwater, sea water and bottled water); you will need access to a solution of soap in ethanol

- Using a measuring cylinder, pour 10 cm^3 of the first water sample into a conical flask
- Fill a burette with the solution of soap in ethanol; record the initial burette reading
- Add 1 cm^3 of the soap solution to the conical flask and shake the flask well; observe whether or not a lather forms
- If no lather forms, add a further 1 cm^3 and shake the flask well; continue this process until a lather forms which lasts for 30 seconds
- Once a lather has formed, record the final burette reading and calculate the volume of soap solution required
- Rinse out your conical flask and repeat steps (a) to (e) for the other water samples
- Record your results in a simple table

Water source	Volume of soap required / cm^3

Which water is the hardest?

Unit 8 – Solubility and Precipitation Reactions

Chemicals needed: water from different sources and a solution of soap in ethanol (around 50 cm³ per group)
Apparatus needed per group: one conical flask, one measuring cylinder (10 cm³), one burette and one funnel
The seawater should be the hardest (need the greatest quantity of soap) and the rainwater should be the softest (need the least quantity of soap)



Extension 7.2: Testing for Temporary and Permanent Hardness in Water

Completed Exercise 7.1? How would you adapt Experiment 7.1 to determine how much of the hardness in each water sample was temporary and how much was permanent?

Take 10 cm³ of water from each source and boil them before adding the soap; then add the soap as in the original experiment; if less soap is required with the boiled sample, some of its hardness is temporary; the bigger the difference, the greater the amount of temporary hardness in the water



Test your knowledge 7.3: Hard and Soft Water

- Which ions are responsible for the hardness of water?
- State two problems caused by hard water
- State one advantage of hard water
- Describe three ways of reducing the hardness of water

- Ca²⁺, Mg²⁺ and Fe²⁺ ions
- Forms limescale when heated, forms scum instead of lather with soap
- It is a good source of minerals for humans
- Distillation, ion exchange, precipitation using sodium carbonate

Unit 8 – Solubility and Precipitation Reactions

Lesson 8 – How much have I learned about solubility and precipitation reactions?



8.1 END-OF-UNIT QUIZ

UNIT 8 – SOLUBILITY AND PRECIPITATION REACTIONS

1. What is a saturated solution?
2. Explain how, in a saturated solution of NaCl, the solid NaCl and the aqueous NaCl form a dynamic equilibrium
3. Describe how the solubility of most solids varies with temperature
4. Describe how the solubility of most gases varies with temperature
5. Use graph paper to draw a solubility curve for NH₄Cl given the following solubility data:

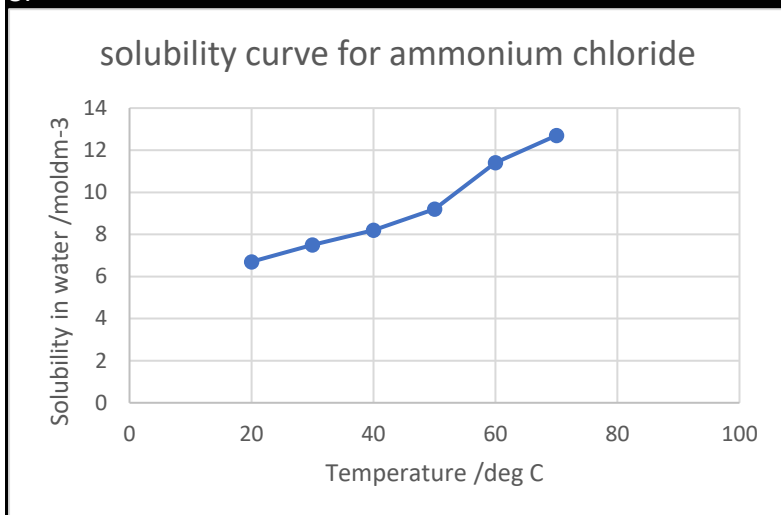
Temperature /°C	Solubility in water /moldm ⁻³
20	6.7
30	7.5
40	8.2
50	9.2
60	11.4
70	12.7

Use your graph to answer the following questions:

- (a) What is the molarity of a saturated solution of NH₄Cl at 25 °C?
 - (b) What mass of NH₄Cl can be dissolved in 50 cm³ of water at 45 °C?
 - (c) What volume of water is needed to dissolve 10 g of NH₄Cl at 55 °C?
 - (d) At what temperature will a 10.0 moldm⁻³ solution of NH₄Cl start to crystallise if it is cooled slowly from 90 °C?
6. Predict whether the following ionic compounds will be soluble in water:
 - (a) copper sulphate
 - (b) barium sulphate
 - (c) sodium carbonate
 - (d) copper carbonate
 - (e) calcium chloride
 7. Describe what you would observe if you mixed the following pairs of solutions together:
 - (a) copper sulphate and sodium hydroxide
 - (b) calcium nitrate and sodium chloride
 - (c) lead nitrate and sulphuric acid
 - (d) calcium chloride and sodium carbonate
 - (e) silver nitrate and hydrochloric acid
 8. Give a simple chemical test which would distinguish between the following solutions:
 - (a) iron (II) sulphate and iron (III) sulphate
 - (b) lead nitrate and zinc nitrate
 - (c) sodium carbonate and sodium sulphate
 9.
 - (a) Which ions are responsible for the hardness of water
 - (b) Give two reasons why hardness of water is a problem
 - (c) Describe three ways in which water can be softened
 - (d) Describe a simple experiment to compare the hardness of different water samples

Unit 8 – Solubility and Precipitation Reactions

1. A solution which contains the maximum amount of dissolved solute which it is possible to dissolve in that quantity of solvent
2. The solid NaCl dissolves and the aqueous NaCl crystallises at equal rates
3. Solubility of solids usually increases with temperature
4. Solubility of gases usually decreases with temperature
- 5.



- (a) $7.1 - 7.2 \text{ moldm}^{-3}$ (b) solubility = $8.5 - 8.7 \text{ moldm}^{-3}$ so mass = 23 – 24 g
(c) solubility = $10.0 - 10.2 \text{ moldm}^{-3}$ so volume = 18 – 19 cm^3 (d) around 54 °C
6. (a) soluble; (b) insoluble; (c) soluble; (d) insoluble; (e) soluble
 7. (a) pale blue precipitate; (b) no reaction; (c) white precipitate; (d) white precipitate; (e) white precipitate
 8. (a) Add NaOH (aq); FeSO_4 gives dark green precipitate, $\text{Fe}_2(\text{SO}_4)_3$ gives orange/brown precipitate
(b) Add HCl (aq); $\text{Pb}(\text{NO}_3)_2$ gives white precipitate; $\text{Zn}(\text{NO}_3)_2$ gives no reaction
(c) Add HCl (aq) and then $\text{BaCl}_2(\text{aq})$; Na_2CO_3 gives no reaction; Na_2SO_4 gives white precipitate
 9. (a) Fe^{2+} , Ca^{2+} , Mg^{2+}
(b) causes limescale when heated, causes scum instead of lather with soap
(c) ion exchange, distillation, precipitation with sodium carbonate
(d) Add soap from a burette to a fixed quantity of different water sample; measure how much is needed to form a lather; the more soap needed, the harder the water